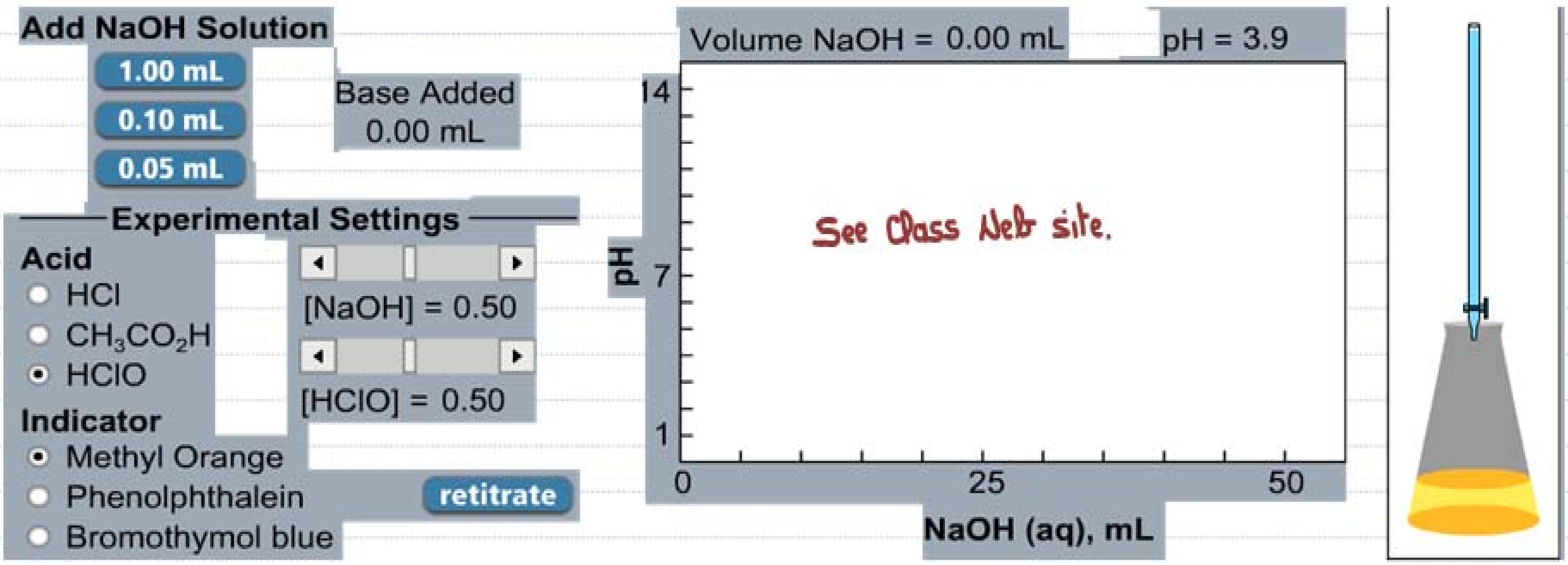
17.3 Acid-Base Titrations

Weak Acid/Strong Base

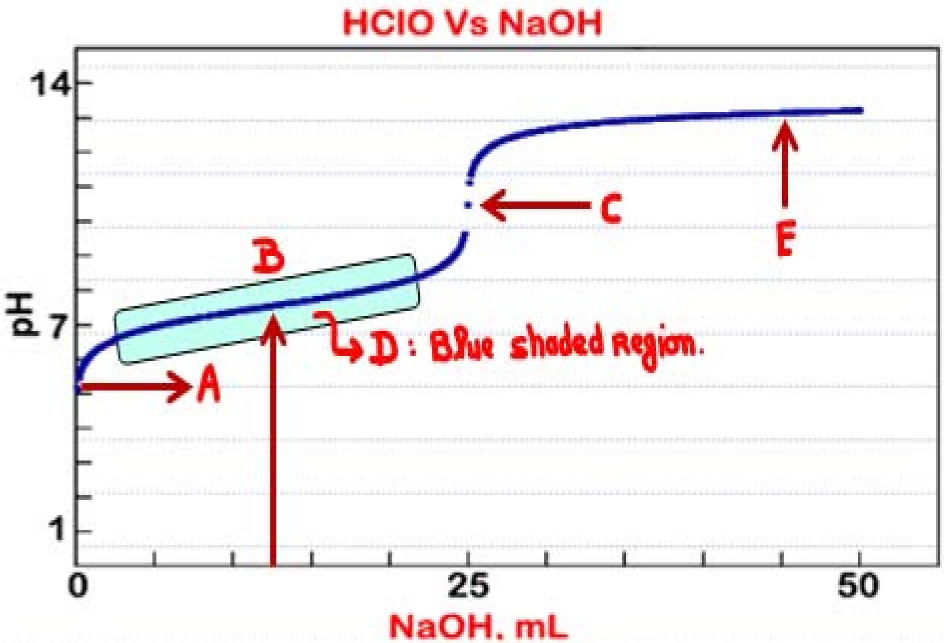
Choose: 0.5M HClO vs 0.5M NaOH



17.3 Acid-Base Titrations

Weak Acid/Strong Base

$HClO(aq) + NaOH(aq) = NaClO(aq) + H_2O(l)$: K_a HClO $= 3.5\times10^{-8}$



D: Buffer Region : $pH = pK_a + \log_{10}\dfrac{[\text{Buffer base}]}{[\text{Buffer acid}]}$

E: $NaOH(aq) = Na^+ + OH^-$

$pOH = -\log_{10}[OH^-]$

$pH = 14 - pOH$

C: Equivalence point, pH depends on hydrolysis

$NaClO(aq) = Na^+ + ClO^-$

Neutral cation. Basic anion.

$ClO^- + H_2O(l) \Leftrightarrow HClO(aq) + OH^-$

$K_b = \dfrac{K_w}{K_a} = \dfrac{1\times10^{-14}}{3.5\times10^{-8}}$

Calculate pOH

$pH = 14 - pOH$

A: $HClO(aq) + H_2O(l) \Leftrightarrow H_3O^+ + ClO^-$

B: [Buffer base] = [Buffer acid]

$[ClO^-] = [HClO]$

pH = pK_a of HClO

18.1 Solubility Equilibria and Ksp

The Solubility Product Constant

Compound	K_{sp} at 25 °C
$PbBr_2$	6.3×10^{-6}
$AgBr$	3.3×10^{-13}
$CaCO_3$	3.8×10^{-9}
$CuCO_3$	2.5×10^{-10}
$NiCO_3$	6.6×10^{-9}
Ag_2CO_3	8.1×10^{-12}
$PbCl_2$	1.7×10^{-5}
$AgCl$	1.8×10^{-10}
BaF_2	1.7×10^{-6}
CaF_2	3.9×10^{-11}
$Cu(OH)_2$	1.6×10^{-19}
$Fe(OH)_3$	6.3×10^{-38}
$Ni(OH)_2$	2.8×10^{-16}
$Zn(OH)_2$	4.5×10^{-17}
$Ca_3(PO_4)_2$	1.0×10^{-25}
$CaSO_4$	2.4×10^{-5}
$PbSO_4$	1.8×10^{-8}

$PbBr_2(s) \rightleftharpoons Pb^{2+} + 2Br^-$

Remember that pure liquids and solids do not appear in an equilibrium expression.

$K = [Pb^{2+}][Br^-]^2$

$\uparrow$

K_{sp} : Solubility Product Constant.

Note that the salts listed are those during Chem III using Solubility Guide lines we would have considered insoluble.

Looking at the Ksp values, these are all reactant-favored equilibria

18.2 Using Ksp in Calculations

Estimating Solubility

Which of the following salts is the **least soluble** in water?

a) AgBr   Ksp = 3.3×10^{-13} @25°C

b) $Cu(OH)_2$ ✓   Ksp = 1.6×10^{-19} @25°C

c) $Ca_3(PO_4)_2$ Ksp = 1.0×10^{-25} @25°C

	AgBr(s)	⇔	Ag^+	+	Br^-
I	Some		0		0
C	-s		s		s
E			s		s

$K_{sp} = [Ag^+][Br^-]$:

$3.3 \times 10^{-13} = (s)(s)$

$s^2 = 3.3 \times 10^{-13}$

$s = \sqrt{3.3 \times 10^{-13}} = 5.47 \times 10^{-7}$

	$Cu(OH)_2$ (s)	⇔	Cu^{2+}	+	2 OH^-
I	Some		0		0
C	-s		s		2s
E			s		2s

$K_{sp} = [Cu^{2+}][OH^-]^2$

$1.6 \times 10^{-19} = (s)(2s)^2$

$4s^3 = 1.6 \times 10^{-19}$

$s^3 = 4.0 \times 10^{-20}$

$s = \sqrt[3]{4.0 \times 10^{-20}} = 3.42 \times 10^{-7}$

	$Ca_3(PO_4)_2$ (s)	⇔	3 Ca^{2+}	+	2 PO_4^{3-}
I	Some		0		0
C	-s		3s		2s
E			3s		2s

$K_{sp} = [Ca^{2+}]^3[PO_4^{3-}]^2$

$1.0 \times 10^{-25} = (3s)^3(2s)^2$

$108s^5 = 1.0 \times 10^{-25}$

$s^5 = 9.3 \times 10^{-28}$

$s = \sqrt[5]{9.3 \times 10^{-28}} = 3.9 \times 10^{-6}$

18.2 Using Ksp in Calculations

Estimating Solubility

General Formula	Example	K_{sp} Expression	K_{sp} as a Function of Molar Solubility (x)	Solubility (x) as a Function of K_{sp}
MY	AgCl	$K_{sp} = [M^+][Y^-]$	$K_{sp} = (x)(x) = x^2$	$x = \sqrt{K_{sp}}$
MY_2	HgI_2	$K_{sp} = [M^{2+}][Y^-]^2$	$K_{sp} = (x)(2x)^2 = 4x^3$	$x = \sqrt[3]{\frac{K_{sp}}{4}}$
MY_3	BiI_3	$K_{sp} = [M^{3+}][Y^-]^3$	$K_{sp} = (x)(3x)^3 = 27x^4$	$x = \sqrt[4]{\frac{K_{sp}}{27}}$
M_2Y_3	$Fe_2(SO_4)_3$	$K_{sp} = [M^{3+}]^2[Y^{2-}]^3$	$K_{sp} = (2x)^2(3x)^3 = 108x^5$	$x = \sqrt[5]{\frac{K_{sp}}{108}}$

Instead of memorizing these simply use the ICE method.

Note that in the ICE table for solubility we use 's' instead of 'x' simply because by solving for s, we have determined the solubility in $mol.L^{-1}$ ... M

18.2 Using Ksp in Calculations

Predicting Whether a Solid Will Precipitate or Dissolve

$$AgCl_{(s)} \rightleftharpoons Ag^+ + Cl^-$$

$$Q = [Ag^+][Cl^-]$$

Compare Q to Ksp